Redox Reactions

What You'll Learn

- You will examine the processes of oxidation and reduction in electron-transfer reactions.
- You will discover how oxidation numbers of elements in compounds are determined and how they relate to electron transfer.
- You will separate redox reactions into their oxidation and reduction processes.
- You will use two different methods to balance oxidation–reduction equations.

Why It's Important

Oxidation and reduction reactions are among the most prevalent in chemistry. From natural phenomena to commercial manufacturing, redox reactions play a major role in your daily life.

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CHEMISTRY Online

Visit the Chemistry Web site at **science.glencoe.com** to find links about the chemistry of oxidation–reduction reactions.

When threatened, the bombardier beetle sprays chemicals from its abdomen that, when combined, undergo an oxidation–reduction reaction. The result is a boiling-hot, foul-smelling "bomb" that allows the beetle to escape predators.

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DISCOVERY LAB

Materials

test tube
iron nail
steel wool or sandpaper
1M copper(II) sulfate ($CuSO_4$)

Observing an Oxidation–Reduction Reaction

Rust is the result of a reaction of iron and oxygen. Iron nails can also react with substances other than oxygen, as you will find out in this experiment.

Safety Precautions

Always wear safety goggles and an apron in the laboratory.

Procedure

1. Use a piece of steel wool to polish the end of an iron nail.
2. Add about 3 mL 1.0M $CuSO_4$ to a test tube. Place the polished end of the nail into the $CuSO_4$ solution. Let stand and observe for about 10 minutes. Record your observations.

Analysis

What is the substance found clinging to the nail? What happened to the color of the copper(II) sulfate solution? Write the balanced chemical equation for the reaction you observed.

Section 20.1 Oxidation and Reduction

Objectives

- **Describe** the processes of oxidation and reduction.
- **Identify** oxidizing and reducing agents.
- **Determine** the oxidation number of an element in a compound.
- **Interpret** redox reactions in terms of change in oxidation state.

Vocabulary

oxidation–reduction reaction
redox reaction
oxidation
reduction
oxidizing agent
reducing agent

In Chapter 10, you learned that a chemical reaction can usually be classified as one of five types—synthesis, decomposition, combustion, single-replacement, or double-replacement. In this chapter, you'll investigate a special characteristic of many of these reactions—the ability of elements to gain or lose electrons when they react with other elements. You experimented with this characteristic when you did the **DISCOVERY LAB**.

Electron Transfer and Redox Reactions

One of the defining characteristics of single-replacement and combustion reactions is that they always involve the transfer of electrons from one atom to another. So do many, but not all, synthesis and decomposition reactions. For example, you studied the synthesis reaction in which sodium and chlorine react to form the ionic compound sodium chloride.

Complete chemical equation: $2Na(s) + Cl_2(g) \longrightarrow 2NaCl(s)$

Net ionic equation: $2Na(s) + Cl_2(g) \longrightarrow 2Na^+ + 2Cl^-$(ions in crystal)

In this reaction, an electron from each of two sodium atoms is transferred to the Cl_2 molecule to form two Cl^- ions. An example of a combustion reaction is the burning of magnesium in air.

Complete chemical equation: $2Mg(s) + O_2(g) \longrightarrow 2MgO(s)$

Net ionic equation: $2Mg(s) + O_2(g) \longrightarrow 2Mg^{2+} + 2O^{2-}$ (ions in crystal)

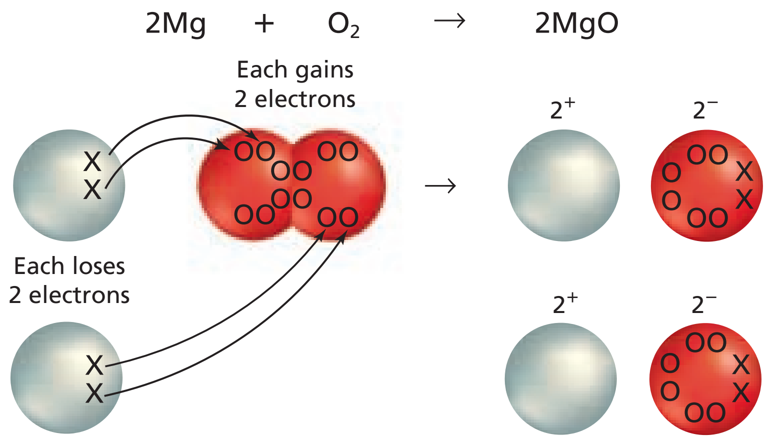


Figure 20-1

The reaction of magnesium and oxygen involves a transfer of electrons from magnesium to oxygen. Therefore, this reaction is an oxidation–reduction reaction. Using the classifications given in Chapter 10, this redox reaction also is classified as a combustion reaction.

When magnesium reacts with oxygen, as illustrated in **Figure 20-1**, each magnesium atom transfers two electrons to each oxygen atom. What is the result of this electron transfer? The two magnesium atoms become Mg^{2+} ions and the two oxygen atoms become O^{2-} ions (oxide ions). If you compare this reaction with the reaction of sodium and chlorine, you will see that they are alike in that both involve the transfer of electrons between atoms. A reaction in which electrons are transferred from one atom to another is called an **oxidation–reduction reaction**. For simplicity, chemists often refer to oxidation–reduction reactions as **redox reactions**.

Now consider the single-replacement reaction in which chlorine in an aqueous solution replaces bromine from an aqueous solution of potassium bromide, which is shown in **Figure 20-2**.

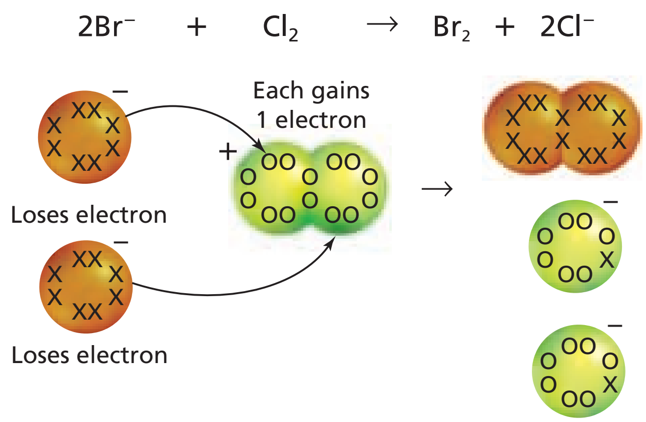
Complete chemical equation: $2KBr(aq) + Cl_2(aq) \longrightarrow 2KCl(aq) + Br_2(aq)$

Net ionic equation: $2Br^-(aq) + Cl_2(aq) \longrightarrow Br_2(aq) + 2Cl^-(aq)$

Note that chlorine "steals" electrons from bromide ions to become chloride ions. When the bromide ions lose their extra electrons, the two bromine atoms form a covalent bond with each other to produce Br_2 molecules. The result of this reaction, the characteristic color of elemental bromine in solution, is shown in **Figure 20-2.** The formation of the covalent bond by sharing of electrons also is an oxidation–reduction reaction.

Figure 20-2

Like the reaction shown in **Figure 20-1,** this reaction of chlorine and bromine in an aqueous solution is also a redox reaction. Here, electrons are transferred from bromide ions to chlorine.



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How do oxidation and reduction differ? Originally, the word *oxidation* referred only to reactions in which a substance combined with oxygen, such as the burning of magnesium in air or the burning of natural gas (methane, CH_4) in air. Today, **oxidation** is defined as the loss of electrons from atoms of a substance. Look again at the net ionic equation for the reaction of sodium and chlorine. Sodium is oxidized because it loses an electron. To state this reaction more clearly,

$$Na \longrightarrow Na^{+} + e^{-}$$

For oxidation to take place, the electrons lost by the substance that is oxidized must be accepted by atoms or ions of another substance. In other words, there must be an accompanying process that involves the gain of electrons. **Reduction** is defined as the gain of electrons by atoms of a substance. Following our sodium chloride example further, the reduction reaction that accompanies the oxidation of sodium is the reduction of chlorine.

$$Cl_2 + 2e^{-} \longrightarrow 2Cl^{-}$$

Can oxidation occur without reduction? By our definitions, oxidation and reduction are complementary processes; oxidation cannot occur unless reduction also occurs.

It is important to recognize and distinguish between oxidation and reduction. The following memory aid may help.

LEO the lion says **GER** or, for short, **LEO GER**

This phrase will help you remember that **L**oss of **E**lectrons is **O**xidation, and **G**ain of **E**lectrons is **R**eduction.

Changes in oxidation number You may recall from previous chapters that the *oxidation number* of an atom in an ionic compound is the number of electrons lost or gained by the atom when it forms ions. For example, look at the following equation for the redox reaction of potassium metal with bromine vapor.

Complete chemical equation: $2K(s) + Br_2(g) \longrightarrow 2KBr(s)$

Net ionic equation: $2K(s) + Br_2(g) \longrightarrow 2K^{+}(s) + 2Br^{-}(s)$

Potassium, a group 1A element that tends to lose one electron in reactions because of its low electronegativity, is assigned an oxidation number of +1. On the other hand, bromine, a group 7A element that tends to gain one electron in reactions because of its high electronegativity, is assigned an oxidation number of −1. In redox terms, you would say that potassium atoms are oxidized from 0 to the +1 state because each loses an electron, and bromine atoms are reduced from 0 to the −1 state because each gains an electron. Can you see why the term reduction is used? When an atom or ion is reduced, the numerical value of its oxidation number is lowered.

Oxidation numbers are tools that scientists use in written chemical equations to help them keep track of the movement of electrons in a redox reaction. Like some of the other tools you have learned about in chemistry, oxidation numbers have a specific notation. Oxidation numbers are written with the positive or negative sign before the number (+3, +2), whereas ionic charge is written with the sign after the number (3+, 2+).

Oxidation number: +3 Ionic charge: 3+

Biology CONNECTION

What do the bacteria *Xenorhabdus luminescens,* fireflies, and many deep-sea fish have in common? These and other organisms emit light. Bioluminescence is the conversion of potential energy in chemical bonds into light during a redox reaction. Bioluminescent light is a cold light that gives off little heat. Depending on the species, bioluminescence is produced by different chemicals and by different means. In fireflies, for example, light results from the oxidation of the molecule luciferin.

Scientists are still unraveling the mystery of bioluminescence. Some luminescent organisms emit light all the time, whereas others emit light only when they are disturbed, such as when churned by ocean waves. Deep-sea fish and some jellyfish appear to be able to control the light they omit, and one species of mushroom is known to emit light of two different colors. Zoologists have also determined that some light-emitting organisms do not produce light themselves, but produce light by harboring bioluminescent bacteria.

Organisms appear to use bioluminescence for different purposes. In the ocean depths, bioluminescence probably aids vision and recognition. Other purposes might include mating and defense against prey.

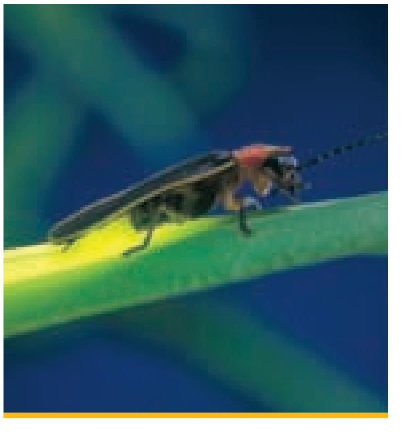

Figure 20-3

A redox reaction involving silver and environmental sulfides deposited tarnish on the discolored cup. Another redox reaction, essentially the reverse of the first, reduces silver ions in the tarnish to silver atoms that can be rinsed away, leaving the shiny silver cup.

Oxidizing and Reducing Agents

Chemists also describe the potassium–bromine reaction in another way by saying that "potassium is oxidized by bromine." This description is useful because it clearly identifies both the substance that is oxidized and the substance that does the oxidizing. The substance that oxidizes another substance by accepting its electrons is called an **oxidizing agent**. This term is another way of saying "the substance that is reduced." The substance that reduces another substance by losing electrons is called a **reducing agent**. A reducing agent supplies electrons to the substance getting reduced (gaining electrons), and is itself oxidized because it loses electrons. By this definition, the reducing agent in the potassium–bromine reaction is potassium, the substance that is oxidized.

oxidized

$$2K(s) + Br_2(g) \longrightarrow 2KBr(s)$$

reduced

A common application of redox chemistry is to remove tarnish from metal objects, such as the silver cups in **Figure 20-3.** The **miniLAB** below describes this tarnish removal technique. Other oxidizing and reducing agents such as those shown in **Figure 20-4** play significant roles in your daily life. For example, when you add chlorine bleach to your laundry to whiten clothes, you are using an aqueous solution of sodium hypochlorite (NaClO), an oxidizing agent. It oxidizes dyes, stains, and other materials that discolor clothes. Hydrogen peroxide (H_2O_2) can be used as an antiseptic because it oxidizes some of the vital biomolecules of germs, or as an agent to lighten hair because it oxidizes the dark pigment of the hair.

miniLAB

Cleaning by Redox

Applying Concepts The tarnish on silver is silver sulfide, which is formed when the silver reacts with sulfide compounds in the environment. In this miniLAB you will use an oxidation–reduction reaction to remove the tarnish from silver or a silver-plated object.

Materials aluminum foil, steel wool, small tarnished silver object, 400-mL beaker (or size large enough to hold the tarnished object), baking soda, table salt, hot plate, beaker tongs

Procedure

1. Buff a piece of aluminum foil lightly with steel wool to remove any oxide coating.
2. Wrap the tarnished object in the aluminum foil, making sure that the tarnished area makes firm contact with the foil.
3. Place the wrapped object in the beaker and add sufficient tap water to cover.
4. Add about 1 spoonful of baking soda and about 1 spoonful of table salt.
5. Set the beaker and contents on a hot plate and heat until the water is nearly boiling. Maintain the heat approximately 15 min until the tarnish disappears.

Analysis

1. Write the equation for the reaction of silver with hydrogen sulfide, yielding silver sulfide and hydrogen.
2. Write the equation for the reaction of the tarnish (silver sulfide) with the aluminum foil, yielding aluminum sulfide and silver.
3. Which metal, aluminum or silver, is more reactive? How do you know this from your results?
4. Why should you not use an aluminum pan to clean silver objects by this method?

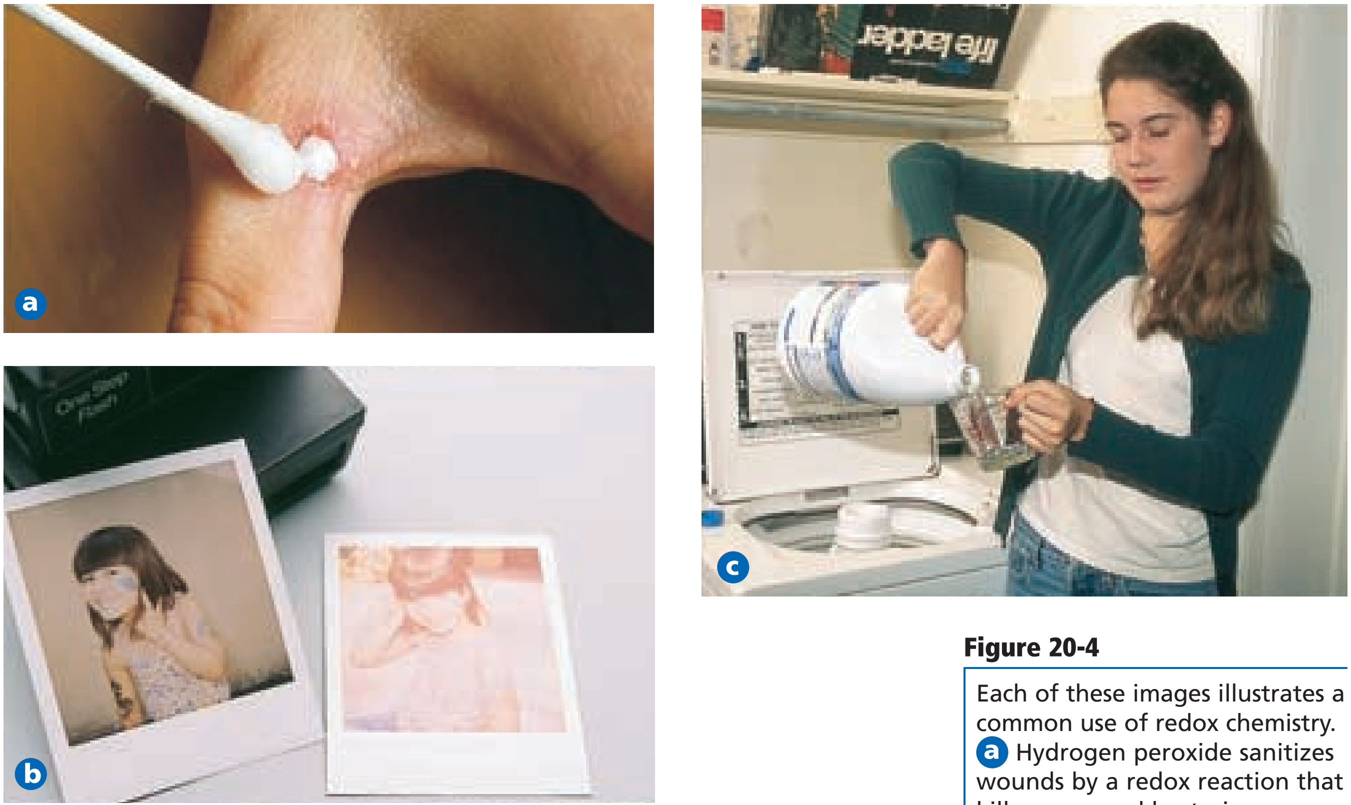

Figure 20-4

Each of these images illustrates a common use of redox chemistry. **a** Hydrogen peroxide sanitizes wounds by a redox reaction that kills germs and bacteria. **b** Photography also uses a series of redox reactions in the image capture and development processes. The feature **How It Works** at the end of this chapter describes how redox reactions are used in photography. **c** Chlorine is a strong oxidizer that is used in chlorine bleach to whiten laundry.

Redox and Electronegativity

The chemistry of oxidation–reduction reactions is not limited to atoms of an element changing to ions or the reverse. Some redox reactions involve changes in molecular substances or polyatomic ions in which atoms are covalently bonded to other atoms. For example, the following equation represents the redox reaction used to manufacture ammonia.

$$N_2(g) + 3H_2(g) \longrightarrow 2NH_3(g)$$

This process involves neither ions nor any obvious transfer of electrons. The reactants and products are all molecular compounds. Still, it is a redox reaction in which nitrogen is the oxidizing agent and hydrogen is the reducing agent.

In situations such as the formation of NH_3 where two atoms share electrons, how is it possible to say that one atom lost electrons and was oxidized while the other atom gained electrons and was reduced? The answer is that you have to know which atom attracts electrons more strongly, or, in other words, which atom is more electronegative. You might find it helpful to review the discussion of electronegativity trends in Chapters 6 and 8. **Table 20-1** gives the specific electronegativities for hydrogen and nitrogen.

Table 20-1

Electronegativity of Hydrogen and Nitrogen	
Hydrogen	**2.20**
Nitrogen	**3.04**

Reduced (partial gain of e^-)

$$N_2(g) + 3H_2(g) \longrightarrow 2NH_3(g)$$

Oxidized (partial loss of e^-)

For the purpose of studying oxidation–reduction reactions, the more electronegative atom (nitrogen) is treated as if it had been reduced by gaining electrons from the other atom. The less electronegative atom (hydrogen) is treated as if it had been oxidized by losing electrons.

CHEMISTRY CD-ROM

Go to the **Chemistry Interactive CD-ROM** to find additional resources for this chapter.

EXAMPLE PROBLEM 20-1

Identifying Oxidation–Reduction Reactions

The following equation represents the redox reaction of aluminum and iron.

$$2Al + 2Fe^{3+} + 3O^{2-} \longrightarrow 2Fe + 2Al^{3+} + 3O^{2-}$$

Identify what is oxidized and what is reduced in this reaction. Identify the oxidizing agent and the reducing agent.

1. Analyze the Problem

You are given the ions in the reaction. Using this information, you must determine the electron transfers that take place. Then you can apply the definitions of oxidizing agent and reducing agent to answer the question.

2. Solve for the Unknown

Identify the oxidation process and the reduction process by evaluating the electron transfer. In this case, aluminum loses three electrons and becomes an aluminum ion in the oxidation process. The iron ion accepts the three electrons lost from aluminum in the reduction process.

$Al \longrightarrow Al^{3+} + 3e^-$ (loss of e^- is oxidation)

$Fe^{3+} + 3e^- \longrightarrow Fe$ (gain of e^- is reduction)

Aluminum is oxidized and therefore is the reducing agent. Iron is reduced and therefore is the oxidizing agent.

3. Evaluate the Answer

In this process aluminum lost electrons and is oxidized, whereas iron gained electrons and is reduced. The definitions of oxidation and reduction and oxidizing agent and reducing agent apply. Note that the oxidation number of oxygen is unchanged in this reaction; therefore, oxygen is not a key factor in this problem.

PRACTICE PROBLEMS

Practice!

For more practice with identifying oxidation-reduction reactions, go to **Supplemental Practice Problems** in Appendix A.

1. Identify each of the following changes as either oxidation or reduction. Recall that e^- is the symbol for an electron.

- **a.** $I_2 + 2e^- \longrightarrow 2I^-$
- **b.** $K \longrightarrow K^+ + e^-$
- **c.** $Fe^{2+} \longrightarrow Fe^{3+} + e^-$
- **d.** $Ag^+ + e^- \longrightarrow Ag$

2. Identify what is oxidized and what is reduced in the following processes.

- **a.** $2Br^- + Cl_2 \longrightarrow Br_2 + 2Cl^-$
- **b.** $Ce + Cu^{2+} \longrightarrow Cu + Ce^{3+}$
- **c.** $2Zn + O_2 \longrightarrow 2ZnO$

3. Identify the oxidizing agent and the reducing agent in each of the following reactions.

- **a.** $Mg + I_2 \longrightarrow MgI_2$
- **b.** $2Na + 2H^+ \longrightarrow 2Na^+ + H_2$
- **c.** $H_2S + Cl_2 \longrightarrow S + 2HCl$

Determining Oxidation Numbers

In order to understand all kinds of redox reactions, you must have a way to determine the oxidation number of the atoms involved in the reaction. Chemists use a set of rules to make this determination easier.

Rules for determining oxidation numbers

1. *The oxidation number of an uncombined atom is zero.* This is true for elements that exist as polyatomic molecules such as O_2, Cl_2, H_2, N_2, S_8.
2. *The oxidation number of a monatomic ion is equal to the charge on the ion.* For example, the oxidation number of a Ca^{2+} ion is +2, and the oxidation number of a Br^- ion is −1.
3. *The oxidation number of the more electronegative atom in a molecule or a complex ion is the same as the charge it would have if it were an ion.* In ammonia (NH_3), for example, nitrogen is more electronegative than hydrogen, meaning that it attracts electrons more strongly than does hydrogen. So nitrogen is assigned an oxidation number of −3, as if it had gained three electrons to complete an octet. In the compound silicon tetrachloride ($SiCl_4$), chlorine is more electronegative than silicon, so each chlorine has an oxidation number of −1 as if it had taken an electron from silicon. The silicon atom is given an oxidation number of +4 as if it had lost electrons to the four chlorine atoms.
4. *The most electronegative element, fluorine, always has an oxidation number of −1 when it is bonded to another element.*
5. *The oxidation number of oxygen in compounds is always −2, except in peroxides, such as hydrogen peroxide (H_2O_2), where it is −1. When it is bonded to fluorine, the only element more electronegative than oxygen, the oxidation number of oxygen is +2.*
6. *The oxidation number of hydrogen in most of its compounds is +1.* The exception to this rule occurs when hydrogen is bonded to less electronegative metals to form hydrides such as LiH, NaH, CaH_2, and AlH_3. In these compounds, hydrogen's oxidation number is −1 because it attracts electrons more strongly than does the metal atom.
7. *The metals of groups 1 and 2 and aluminum in group 3A form compounds in which the metal atom always has a positive oxidation number equal to the number of its valence electrons (+1, +2, and +3, respectively).*
8. *The sum of the oxidation numbers in a neutral compound is zero.* Notice how the oxidation numbers add up to zero in the following examples.

$$\underset{NaCl}{(+1) + (-1) = 0} \qquad \underset{CaBr_2}{(+2) + 2(-1) = 0} \qquad \underset{H_2S}{2(+1) + (-2) = 0} \qquad \underset{CCl_4}{(+4) + 4(-1) = 0}$$

9. *The sum of the oxidation numbers of the atoms in a polyatomic ion is equal to the charge on the ion.* The following examples illustrate.

$$\underset{NH_4^{+}}{(-3) + 4(+1) = +1} \qquad \underset{SO_3^{2-}}{(+4) + 3(-2) = -2}$$

Many elements other than those specified in the rules above, including most of the transition metals, metalloids, and nonmetals, can be found with different oxidation numbers in different compounds. For example, the two copper compounds and the two chromium compounds shown in **Figure 20-5 a** and **b,** respectively, have different oxidation numbers.

Careers Using Chemistry

Photochemical Etching Artist

Does the idea of using photographic equipment and chemicals as art appeal to you? Then you might want to investigate a career in photochemical etching.

Photochemical etching was first used to produce precise machine parts, such as microchips and other electronic devices. Now artists use it to create metal art. In this process, ultraviolet light is used to transfer a pattern onto a piece of metal. Then chemicals are applied to remove certain areas in the pattern, creating an intricate design on the metal.

Figure 20-5

The difference in oxidation number gives compounds different chemical and physical properties.
a The oxidation number of copper in copper(I) oxide (orange color) is +1, whereas the oxidation number of copper in copper(II) oxide (black color) is +2.
b Likewise, the oxidation number of chromium in chromium(II) chloride tetrahydrate (blue solution) differs from that in chromium(III) chloride hexahydrate (green solution).

EXAMPLE PROBLEM 20-2

Determining Oxidation Numbers

Determine the oxidation number of each element in the following compounds or ions.

a. $KClO_3$ (potassium chlorate)

b. SO_3^{2-} (sulfite ion)

1. Analyze the Problem

From the rules for determining oxidation numbers, you are given the oxidation number of oxygen (−2) and potassium (group 1 metal, +1). You are also given the overall charge of the compound or ion. Using this information and applying the rules, you can determine the oxidation numbers of chlorine and sulfur.

2. Solve for the Unknown

Strategy

Assign the known oxidation numbers to their elements, set the sum of all oxidation numbers to zero or to the ion charge, and solve for the unknown oxidation number. (Let $n_{element}$ = oxidation number of the element in question.)

Solution

a. Potassium chlorate is a neutral salt, so oxidation numbers must add up to zero. According to rule 5, the oxidation number of oxygen in compounds is −2. Rule 7 states that Group 1 metals have a +1 oxidation number in compounds.

$$\overset{(+1) + (n_{Cl}) + 3(-2) = 0}{KClO_3}$$

$$1 + n_{Cl} + (-6) = 0$$

$$n_{Cl} = +5$$

b. Sulfite ion has a charge of 2−, so oxidation numbers must add up to −2. According to rule 5, the oxidation number of oxygen in compounds is −2.

$$\overset{(n_S) + 3(-2) = -2}{SO_3^{2-}}$$

$$n_S + (-6) = -2$$

$$n_S = +4$$

3. Evaluate the Answer

The rules for determining oxidation numbers have been correctly applied. All of the oxidation numbers in each substance add up to the proper value.

Among other applications, potassium chlorate is used in explosives, fireworks, and matches.

PRACTICE PROBLEMS

4. Determine the oxidation number of the boldface element in the following formulas for compounds.

   a. Na**Cl**O_4 **b.** Al**P**O_4 **c.** H**N**O_2

5. Determine the oxidation number of the boldface element in the following formulas for ions.

   a. **N**H_4^+ **b.** **As**O_4^{3-} **c.** **Cr**O_4^{2-}

For more practice with determining oxidation numbers, go to **Supplemental Practice Problems** in Appendix A.

Oxidation Number in Redox Reactions

Having studied oxidation numbers, you are now able to relate oxidation–reduction reactions to changes in oxidation number. Look again at the equation for a reaction that you saw at the beginning of this section, the replacement of bromine in aqueous KBr by Cl_2.

$$2KBr(aq) + Cl_2(aq) \longrightarrow 2KCl(aq) + Br_2(aq)$$

To see how oxidation numbers change, start by assigning numbers, shown in **Table 20-2,** to all elements in the balanced equation. Then review the changes as shown in the accompanying diagram.

Table 20-2

Oxidation Number Assignment		
Element	**Oxidation number**	**Rule**
K in KBr	+1	7
Br in KBr	−1	8
Cl in Cl_2	0	1
K in KCl	+1	7
Cl in KCl	−1	8
Br in Br_2	0	1

Change: −1
reduction

Change: +1
oxidation

$$\overset{+1\ -1}{2KBr(aq)} + \overset{0}{Cl_2(aq)} \longrightarrow \overset{+1\ -1}{2KCl(aq)} + \overset{0}{Br_2(aq)}$$

No change in
oxidation number

As you can see, the oxidation number of bromine changed from −1 to 0, an increase of 1. At the same time, the oxidation number of chlorine changed from 0 to −1, a decrease of 1. Therefore, chlorine is reduced and bromine is oxidized. All redox reactions follow the same pattern. When an atom is oxidized, its oxidation number increases. When an atom is reduced, its oxidation number decreases. Note that there is no change in the oxidation number of potassium. The potassium ion takes no part in the reaction and is called a spectator ion. How would the reaction differ if you used zinc bromide ($ZnBr_2$) instead of potassium bromide?

CHEMISTRY CD-ROM

Go to the **Chemistry Interactive CD-ROM** to find additional resources for this chapter.

Section 20.1 Assessment

6. Describe the processes of oxidation and reduction.

7. Explain the roles of oxidizing agents and reducing agents in a redox reaction. How is each changed in the reaction?

8. Determine the oxidation number of the boldface element in these compounds.

a. **Sb**$_2O_5$ **c.** Ca**N**$_2$
b. H**N**O_3 **d.** Cu**W**O_4 (copper(II) tungstate)

9. Determine the oxidation number of the boldface element in these ions.

a. **I**O_4^- **c.** **B**$_4O_7^{2-}$
b. **Mn**O_4^- **d.** **N**H_2^-

10. **Thinking Critically** Write the equation for the reaction of iron metal with hydrobromic acid to form iron(III) bromide and hydrogen gas. Determine which element is reduced and which element is oxidized in this reaction.

11. **Making Predictions** Alkali metals are strong reducing agents. Would you predict that their reducing ability would increase or decrease as you move down the family from sodium to francium? Give reasons for your prediction.

Section 20.2 Balancing Redox Equations

Objectives

- **Relate** changes in oxidation numbers to the transfer of electrons.
- **Use** changes in oxidation number to balance redox equations.
- **Balance** net ionic redox equations by the oxidation-number method.

Vocabulary

oxidation-number method

You already know that chemical equations are written to represent chemical reactions by showing what substances react and what products are formed. You also know that chemical equations must be balanced to show the correct quantities of reactants and products. Equations for oxidation–reduction reactions are no different. In this section, you'll learn a specific method to balance redox equations.

The Oxidation-Number Method

Many equations for redox reactions are easy to balance. For example, try to balance the following unbalanced equation for the redox reaction that occurs when potassium chlorate is heated and decomposes to produce potassium chloride and oxygen gas.

$$\text{Unbalanced equation: } KClO_3(s) \longrightarrow KCl(s) + O_2(g)$$

Did you get the correct result shown below?

$$\text{Balanced equation: } 2KClO_3(s) \longrightarrow 2KCl(s) + 3O_2(g)$$

Equations for other redox reactions are not as easy to balance. Study the following unbalanced equation for the reaction that occurs when copper metal is placed in concentrated nitric acid. This reaction is shown in **Figure 20-6.** The brown gas you see is nitrogen dioxide (NO_2), produced by the reduction of nitrate ions (NO_3^-), and the blue solution is the result of the oxidation of copper to copper(II) ions.

$$Cu(s) + HNO_3(aq) \longrightarrow Cu(NO_3)_2(aq) + NO_2(g) + H_2O(l)$$

Figure 20-6

Some chemical equations for redox reactions, such as this reaction between copper and nitric acid, can be difficult to balance because one or more elements may appear several times on both sides of the equation. The oxidation-number method makes it easier to balance the equation for this redox reaction.

Note that oxygen appears in only one reactant, HNO_3, but in all three products. Nitrogen appears in HNO_3 and in two of the products. Redox equations such as this one, in which the same element appears in several reactants and products, can be hard to balance by the conventional method. For this reason, you need to learn a different balancing technique for redox equations that is based on the fact that the number of electrons transferred from atoms must equal the number of electrons accepted by other atoms.

As you know, when an atom loses electrons, its oxidation number increases; when an atom gains electrons, its oxidation number decreases. Therefore, the total increase in oxidation numbers (oxidation) must equal the total decrease in oxidation numbers (reduction) of the atoms involved in the reaction. The balancing technique called the **oxidation-number method** is based on these principles. The **CHEMLAB** at the end of this chapter gives you the opportunity to perform and balance the copper–nitric acid redox reaction.

Steps for balancing redox equations by the oxidation-number method

1. Assign oxidation numbers to all atoms in the equation.
2. Identify the atoms that are oxidized and the atoms that are reduced.
3. Determine the change in oxidation number for the atoms that are oxidized and for the atoms that are reduced.
4. Make the change in oxidation numbers equal in magnitude by adjusting coefficients in the equation.

5. If necessary, use the conventional method to balance the remainder of the equation.

You can see these steps clearly by following Example Problem 20-3.

EXAMPLE PROBLEM 20-3

Balancing a Redox Equation by the Oxidation-Number Method

Balance the redox equation shown here for the reaction that produces copper nitrate.

$Cu + HNO_3 \longrightarrow Cu(NO_3)_2 + NO_2 + H_2O$

1. Analyze the Problem

You are given the formulas for the reactants and products, and you have the rules for determining oxidation number. You also know that the increase in oxidation number of the oxidized atoms must equal the decrease in oxidation number of the reduced atoms. With this information you can adjust the coefficients to balance the equation.

2. Solve for the Unknown

Step 1. Apply the appropriate rules on page 641 to assign oxidation numbers to all atoms in the equation.

0   +1 +5 −2   +2 +5 −2   +4 −2   +1 −2

$Cu + HNO_3 \longrightarrow Cu(NO_3)_2 + NO_2 + H_2O$

Step 2. Identify which atoms are oxidized and which are reduced.

0   +1 +5 −2   +2 +5 −2   +4 −2   +1 −2

$Cu + HNO_3 \longrightarrow Cu(NO_3)_2 + NO_2 + H_2O$

↑   ↑   ↑   ↑

The oxidation number of copper increases from 0 to +2 as it is oxidized in the reaction. The oxidation number of nitrogen decreases from +5 to +4 as it is reduced in the formation of NO_2. The oxidation numbers of hydrogen and oxygen do not change. Note that the nitrogen atoms remain unchanged in the nitrate ion (NO_3^-), which appears on both sides of the equation. It is neither oxidized nor reduced.

Step 3. Draw a line connecting the atoms involved in oxidation and another line connecting the atoms involved in reduction. Write the net change in oxidation number above or below each line.

+2

$Cu + HNO_3 \longrightarrow Cu(NO_3)_2 + NO_2 + H_2O$

−1

Step 4. Make the changes in oxidation number equal in magnitude by placing the appropriate coefficients in front of the formulas in the equation. Because the oxidation number for nitrogen is −1, you must add a coefficient 2 to balance. This coefficient applies to HNO_3 on the left and NO_2 on the right.

+2

$Cu + 2HNO_3 \longrightarrow Cu(NO_3)_2 + 2NO_2 + H_2O$

2(−1) = −2

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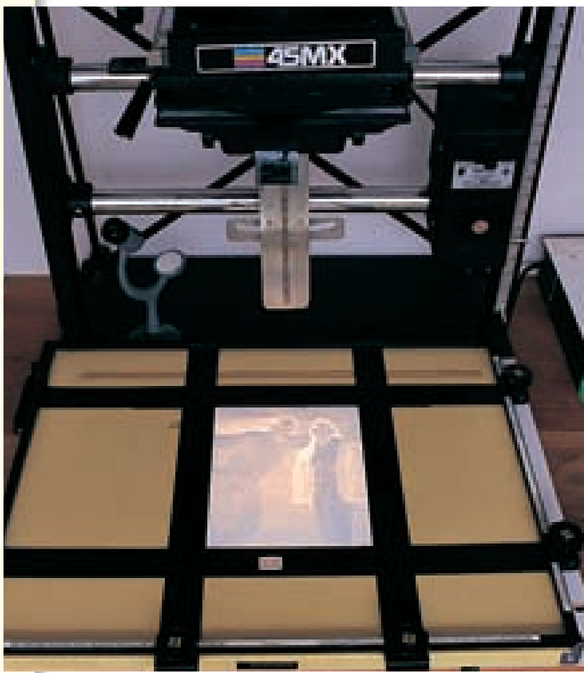


Copper nitrate ($Cu(NO_3)_2$) is used in light-sensitive paper.

Step 5. Use the conventional method to balance the number of atoms and the charges in the remainder of the equation. This equation shows no charged species, so charge balance need not be considered here.

$Cu + 2HNO_3 \longrightarrow Cu(NO_3)_2 + 2NO_2 + H_2O$

The coefficient of HNO_3 must be increased from 2 to 4 to balance the four nitrogen atoms in the products.

$Cu + 4HNO_3 \longrightarrow Cu(NO_3)_2 + 2NO_2 + H_2O$

Add a coefficient of 2 to H_2O to balance the four hydrogen atoms on the left.

$Cu(s) + 4HNO_3(aq) \longrightarrow Cu(NO_3)_2(aq) + 2NO_2(g) + 2H_2O(l)$

3. Evaluate the Answer

The number of atoms of each element is equal on both sides of the equation. No subscripts have been changed.

PRACTICE PROBLEMS

Use the oxidation-number method to balance these redox equations.

12. $HCl + HNO_3 \longrightarrow HOCl + NO + H_2O$

13. $SnCl_4 + Fe \longrightarrow SnCl_2 + FeCl_3$

14. $NH_3(g) + NO_2(g) \longrightarrow N_2(g) + H_2O(l)$

Practice!

For more practice balancing redox reactions, go to **Supplemental Practice Problems** in Appendix A.

Balancing Net Ionic Redox Equations

Sometimes chemists prefer to express redox reactions in the simplest possible terms—as an equation showing only the oxidation and reduction processes and nothing else. Look again at the balanced equation for the oxidation of copper by nitric acid.

$$Cu(s) + 4HNO_3(aq) \longrightarrow Cu(NO_3)_2(aq) + 2NO_2(g) + 2H_2O(l)$$

Note that the reaction takes place in aqueous solution, so HNO_3, which is a strong acid, will be ionized. Likewise, copper(II) nitrate ($Cu(NO_3)_2$) will be dissociated into ions. Therefore, the equation can also be written in ionic form.

$$Cu(s) + 4H^+(aq) + 4NO_3^-(aq) \longrightarrow$$
$$Cu^{2+}(aq) + 2NO_3^-(aq) + 2NO_2(g) + 2H_2O(l)$$

You can see that there are four nitrate ions among the reactants, but only two of them undergo change to form two nitrogen dioxide molecules. What is the role of the other nitrate ions? The other two are only spectator ions and can be eliminated from the equation. To simplify things when writing redox equations in ionic form, chemists usually indicate hydrogen ions by H^+ with the understanding that they exist in hydrated form as hydronium ions (H_3O^+). The equation can then be rewritten showing only the substances that undergo change.

$$Cu(s) + 4H^+(aq) + 2NO_3^-(aq) \longrightarrow Cu^{2+}(aq) + 2NO_2(g) + 2H_2O(l)$$

Now look at the equation in unbalanced form.

$$Cu(s) + H^+(aq) + NO_3^-(aq) \longrightarrow Cu^{2+}(aq) + NO_2(g) + H_2O(l)$$

You might even see this same reaction expressed in a way that shows only the substances that are oxidized and reduced.

$$Cu(s) + NO_3^-(aq) \longrightarrow Cu^{2+}(aq) + NO_2(g) \text{ (in acid solution)}$$

In this case, the hydrogen ion and the water molecule are eliminated because neither is oxidized nor reduced. The only additional information needed is that the reaction takes place in acid solution. In acid solution, hydrogen ions (H^+) and water molecules are abundant and free to participate in redox reactions as either reactants or products. Some redox reactions can occur only in basic solution. When you balance equations for these reactions, you may add hydroxide ions (OH^-) and water molecules to either side of the equation. Basic solutions have an abundance of OH^- ions instead of H_3O^+ ions.

If you try to balance the equation given above, you will see that it appears impossible. Net ionic equations can still be balanced, though, by applying the oxidation-number method, as you will see in Example Problem 20-4. The **problem-solving LAB** below shows how the oxidation-number method can be used for a real-world application.

problem-solving LAB

How does redox lift a space shuttle?

Using Numbers The space shuttle gains nearly 72% of its lift from its solid rocket boosters (SRBs) during the first two minutes of launch. The two pencil-shaped SRB tanks are attached to both sides of the liquid hydrogen and oxygen fuel tank. Each SRB contains 495 000 kg of an explosive mixture of ammonium perchlorate and aluminum. The unbalanced equation for the reaction is given below.

$$NH_4ClO_4(s) + Al(s) \longrightarrow Al_2O_3(g) + HCl(g) + N_2(g) + H_2O(g)$$

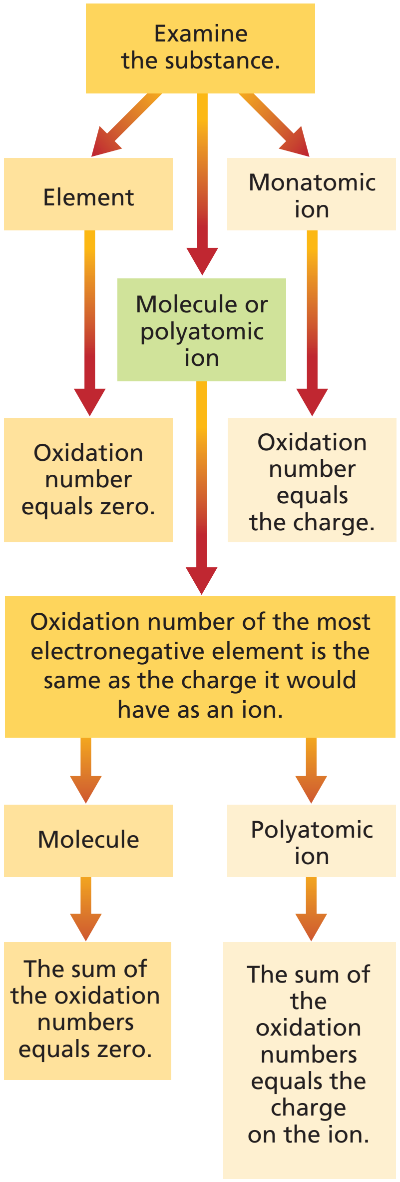
Use the oxidation-number method to balance this redox equation.

Analysis

The SRB reaction can be balanced using the rules for assigning oxidation numbers. In this way, you can ensure that the oxidation reaction balances the reduction reaction. Write the oxidation number above each element in the equation. Which elements are reduced and which are oxidized? What coefficients are required to balance the reaction?

Thinking Critically

What are some of the benefits of using SRBs for the first two minutes of launch?



Use this flow chart to review the rules for assigning oxidation numbers listed on page 641.

EXAMPLE PROBLEM 20-4

Balancing a Net Ionic Redox Equation

Use the oxidation-number method to balance this net ionic redox equation for the reaction between the perchlorate ion and the bromide ion in acid solution.

$ClO_4^-(aq) + Br^-(aq) \rightarrow Cl^-(aq) + Br_2(g)$ (in acid solution)

1. Analyze the Problem

You are given the formulas for the reactants and products, and you have the rules for determining oxidation number. You also know that the increase in oxidation number of the oxidized atoms must equal the decrease in oxidation number of the reduced atoms. You are told that the reaction takes place under acidic conditions. With this information you can adjust the coefficients to balance the equation.

2. Solve for the Unknown

Step 1. Apply rules to assign oxidation numbers to all atoms in the equation using the flow chart on the left.

$$\overset{+7\,-2}{ClO_4^-}(aq) + \overset{-1}{Br^-}(aq) \rightarrow \overset{-1}{Cl^-}(aq) + \overset{0}{Br_2}(g) \text{ (in acid solution)}$$

Step 2. Identify which atoms are oxidized and which are reduced.

$$\overset{+7\,-2}{ClO_4^-}(aq) + \overset{-1}{Br^-}(aq) \rightarrow \overset{-1}{Cl^-}(aq) + \overset{0}{Br_2}(g) \text{ (in acid solution)}$$

The oxidation number of bromine increases from −1 to 0 as it is oxidized. The oxidation number of chlorine decreases from +7 to −1 as it is reduced. Note that no oxygen atoms appear in the products. This deficiency will be fixed shortly.

Step 3. Draw a line connecting the atoms involved in oxidation and another line connecting the atoms involved in reduction. Write the net change in oxidation number above or below each line.

+1

$$ClO_4^-(aq) + Br^-(aq) \rightarrow Cl^-(aq) + Br_2(g) \text{ (in acid solution)}$$

−8

Step 4. Make the changes in oxidation number equal in magnitude by placing the appropriate coefficients in front of the formulas in the equation.

8(+1)=+8

$$ClO_4^-(aq) + 8Br^-(aq) \rightarrow Cl^-(aq) + 4Br_2(g) \text{ (in acid solution)}$$

−8

Note that $4Br_2$ represents eight bromine atoms to balance the $8Br^-$ on the left side.

Step 5. Add enough hydrogen ions and water molecules to the equation to balance the oxygen atoms on both sides. This is the reason that you must know the conditions of the reaction.

$$ClO_4^-(aq) + 8Br^-(aq) + 8H^+(aq) \rightarrow Cl^-(aq) + 4Br_2(g) + 4H_2O(l)$$

CONTENTS

Four oxygen atoms are present in one ClO_4^- ion, enough to form four H_2O molecules. Therefore, a total of $8H^+$ ions will also be needed to form the four water molecules.

3. Evaluate the Answer

Review the balanced equation, counting the number of atoms to be sure there are equal numbers of each element. As with any ionic equation, you must also check to see if the net charge on the right equals the net charge on the left. Check to be sure that you did not change any subscripts in the equation.

PRACTICE PROBLEMS

Use the oxidation-number method to balance the following net ionic redox equations.

15. $H_2S(g) + NO_3^-(aq) \longrightarrow S(s) + NO(g)$ (in acid solution)

16. $Cr_2O_7^{2-}(aq) + I^-(aq) \longrightarrow Cr^{3+}(aq) + I_2(s)$ (in acid solution)

17. $I^-(aq) + MnO_4^-(aq) \longrightarrow I_2(s) + MnO_2(s)$ (in basic solution) Hint: Hydroxide ions will appear on the right, and water molecules on the left.

Practice!

For more practice balancing redox reactions, go to **Supplemental Practice Problems** in Appendix A.

The oxidation-number method is convenient for balancing most redox equations, but you will see in the next section that there are occasions when you must balance the net ionic charge on both sides of the equation in addition to balancing the atoms.

Section 20.2 Assessment

18. A reactant in an oxidation–reduction reaction loses four electrons when it is oxidized. How many electrons must be gained by the reactant that is reduced?

19. Why is it important to know the conditions under which an aqueous redox reaction takes place in order to balance the ionic equation for the reaction?

20. Balance these equations for redox reactions by using the oxidation-number method.

a. $HClO_3(aq) \longrightarrow ClO_2(g) + HClO_4(aq) + H_2O(l)$

b. $H_2O_2(aq) + H_2SO_4(aq) + FeSO_4(aq) \longrightarrow Fe_2(SO_4)_3(aq) + H_2O(l)$
(Hint: Review rule 5 to learn how to determine the oxidation numbers in H_2O_2.)

c. $H_2SeO_3(aq) + HClO_3(aq) \longrightarrow H_2SeO_4(aq) + Cl_2(g) + H_2O(l)$

21. Balance these net ionic equations for redox reactions.

a. $Cr_2O_7^{2-}(aq) + Fe^{2+}(aq) \longrightarrow Cr^{3+}(aq) + Fe^{3+}(aq)$ (in acid solution)

b. $Zn(s) + V_2O_5(aq) \longrightarrow Zn^{2+}(aq) + V_2O_4(aq)$ (in acid solution)

c. $N_2O(g) + ClO^-(aq) \longrightarrow Cl^-(aq) + NO_2^-(aq)$ (in basic solution)

22. Thinking Critically Explain how changes in oxidation number are related to the electrons transferred in a redox reaction. How are the changes related to the processes of oxidation and reduction?

23. Applying Concepts The processes used to remove metals from their ores usually involve a reduction process. For example, mercury may be obtained by roasting the ore mercury(II) sulfide with calcium oxide to produce mercury metal, calcium sulfide, and calcium sulfate. Write and balance the redox equation for this process.

Section 20.3

Half-Reactions

Objectives

- **Recognize** the interdependence of oxidation and reduction processes.
- **Derive** oxidation and reduction half-reactions from a redox equation.
- **Balance** redox equations by the half-reaction method.

Vocabulary

species
half-reaction

Throughout this chapter, you have read about oxidation–reduction reactions. You know that redox reactions involve the loss and gain of electrons. Thus, the "pairing" or complementary nature of redox reactions is probably apparent to you. So, let's consider the two halves of redox reactions.

Identifying Half-Reactions

As you know, oxidation–reduction reactions can involve molecules, ions, free atoms, or combinations of all three. To make it easier to discuss redox reactions without constantly specifying the kind of particle involved, chemists use the term species. In chemistry, a **species** is any kind of chemical unit involved in a process. For example, a solution of sugar in water contains two major species. In the equilibrium equation $NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$, there are four species: the two molecules NH_3 and H_2O and the two ions NH_4^+ and OH^-.

Oxidation–reduction reactions take place whenever a species that can give up electrons (reducing agent) comes in contact with another species that can accept them (oxidizing agent). In a sense, the species being reduced doesn't "care" where the electrons are coming from as long as they are readily available. Likewise, the species undergoing oxidation doesn't "care" where its electrons go as long as a willing receiver is available. The cartoon electron donor/acceptor machine shown in **Figure 20-7** illustrates this principle.

For example, active metals can be oxidized by a variety of oxidizing agents. When iron rusts, it is oxidized by oxygen. You can also say that iron reduces oxygen because iron acts as a reducing agent.

$$4Fe + 3O_2 \longrightarrow 2Fe_2O_3$$

But iron can reduce many other oxidizing agents, including chlorine.

$$2Fe + 3Cl_2 \longrightarrow 2FeCl_3$$

In this reaction, each iron atom is oxidized by losing three electrons to become an Fe^{3+} ion. This is the oxidation half of the oxidation–reduction reaction. Recall that e^- represents one electron.

$$Fe \longrightarrow Fe^{3+} + 3e^-$$

At the same time, each chlorine atom in Cl_2 is reduced by accepting one electron to become a Cl^- ion. This is the reduction half of the oxidation-reduction reaction.

$$Cl_2 + 2e^- \longrightarrow 2Cl^-$$

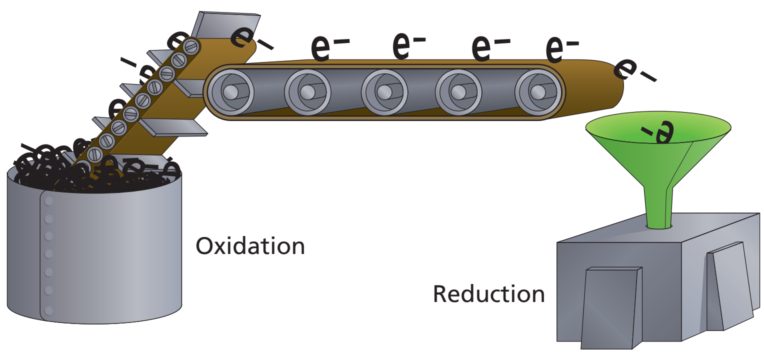


Figure 20-7

This imaginary electron machine illustrates the point that the electrons that are involved in redox reactions can come from any source, as long as they are available to transfer. It also points out that reduction, the complimentary process of oxidation, requires something to accept the electrons. These two processes can be identified and separated to help understand redox reactions and balance redox equations.

Table 20-3

Various Oxidation–Reduction Reactions in which Iron Is Oxidized		
Overall reaction (unbalanced)	**Oxidation half-reaction**	**Reduction half-reaction**
$Fe + O_2 \rightarrow Fe_2O_3$	$Fe \rightarrow Fe^{3+} + 3e^-$	$O_2 + 4e^- \rightarrow 2O^{2-}$
$Fe + Cl_2 \rightarrow FeCl_3$	$Fe \rightarrow Fe^{3+} + 3e^-$	$Cl_2 + 2e^- \rightarrow 2Cl^-$
$Fe + F_2 \rightarrow FeF_3$	$Fe \rightarrow Fe^{3+} + 3e^-$	$F_2 + 2e^- \rightarrow 2F^-$
$Fe + HBr \rightarrow H_2 + FeBr_3$	$Fe \rightarrow Fe^{3+} + 3e^-$	$2H^+ + 2e^- \rightarrow H_2$
$Fe + AgNO_3 \rightarrow Ag + Fe(NO_3)_3$	$Fe \rightarrow Fe^{3+} + 3e^-$	$Ag^+ + e^- \rightarrow Ag$
$Fe + CuSO_4 \rightarrow Cu + Fe_2(SO_4)_3$	$Fe \rightarrow Fe^{3+} + 3e^-$	$Cu^{2+} + 2e^- \rightarrow Cu$

Equations such as these represent half-reactions. A **half-reaction** is one of the two parts of a redox reaction—the oxidation half alone or the reduction half alone. **Table 20-3** shows a variety of reduction half-reactions that can accept electrons from iron, oxidizing it from Fe to Fe^{3+}.

Balancing Redox Equations by Half-Reactions

You will learn more about the importance of half-reactions when you study electrochemistry in Chapter 21. For now, however, you can learn to use half-reactions to balance a redox equation. First, look at an unbalanced equation taken from **Table 20-3** to see how to separate a redox equation into half-reactions. For example, the following unbalanced equation represents the reaction that occurs when you put an iron nail into a solution of copper(II) sulfate, as shown in **Figure 20-8.** Iron atoms are oxidized as they lose electrons to the copper(II) ions.

$$Fe(s) + CuSO_4(aq) \rightarrow Cu(s) + Fe_2(SO_4)_3(aq)$$

Steps for balancing by half-reactions

1. Write the net ionic equation for the reaction, omitting spectator ions.

$Fe + Cu^{2+} + \cancel{SO_4^{2-}} \rightarrow Cu + 2Fe^{3+} + \cancel{3SO_4^{2-}}$

$Fe + Cu^{2+} \rightarrow Cu + 2Fe^{3+}$

2. Write the oxidation and reduction half-reactions for the net ionic equation.

$Fe \rightarrow 2Fe^{3+} + 6e^-$

$Cu^{2+} + 2e^- \rightarrow Cu$

3. Balance the atoms and charges in each half-reaction.

$2Fe \rightarrow 2Fe^{3+} + 6e^-$

$Cu^{2+} + 2e^- \rightarrow Cu$

4. Adjust the coefficients so that the number of electrons lost in oxidation equals the number of electrons gained in reduction.

$2Fe \rightarrow 2Fe^{3+} + \cancel{6e^-}$

$3Cu^{2+} + \cancel{6e^-} \rightarrow 3Cu$

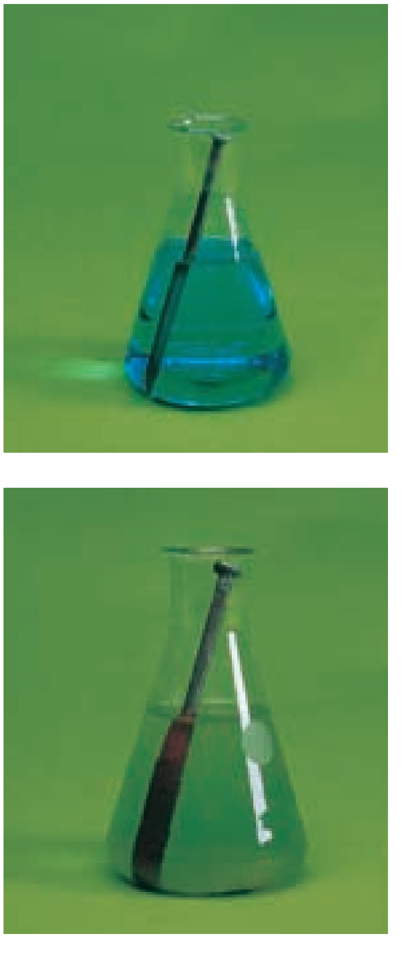
5. Add the balanced half-reactions and return spectator ions.

$2Fe + 3Cu^{2+} \rightarrow 3Cu + 2Fe^{3+}$

$2Fe(s) + 3CuSO_4(aq) \rightarrow 3Cu(s) + Fe_2(SO_4)_3(aq)$

Figure 20-8

As a result of this redox reaction between iron and copper sulfate solution, solid copper metal is deposited on the iron nail. To balance the equation given in the text for this reaction you could use the method of half-reactions.



EXAMPLE PROBLEM 20-5

Balancing a Redox Equation by Half-Reactions

The permanganate ion (MnO_4^-) is widely used in the chemistry laboratory as a strong oxidizing agent. Permanganate is usually sold as potassium salt $KMnO_4$. When sulfur dioxide gas is bubbled into an acidic solution of potassium permanganate, it reacts to form sulfate ions. Balance the redox equation for this reaction using half-reactions.

$$KMnO_4(aq) + SO_2(g) \rightarrow MnSO_4(aq) + K_2SO_4(aq)$$

1. Analyze the Problem

You are given the skeleton equation for the reaction of permanganate and sulfur dioxide. You also know that the reaction takes place in an acid solution. With this information, the rules for determining oxidation numbers, and the steps for balancing by half-reactions, you can write a complete balanced equation.

2. Solve for the Unknown

Step 1. Write the net ionic equation for the reaction. You can eliminate coefficients, spectator ions, and state symbols.

$$MnO_4^- + SO_2 \rightarrow Mn^{2+} + SO_4^{2-}$$

Step 2. Write the oxidation and reduction half-reactions for the net ionic equation, including oxidation numbers. Recall the rules for assigning oxidation numbers from page 641.

$$\overset{+4}{S}O_2 \rightarrow \overset{+6}{S}O_4^{2-} + 2e^- \text{ (oxidation)}$$

$$\overset{+7}{Mn}O_4^- + 5e^- \rightarrow \overset{+2}{Mn}{}^{2+} \text{ (reduction)}$$

Potassium permanganate ($KMnO_4$) is used to disinfect decorative fish ponds and the water in fish hatcheries.

Step 3. Balance the atoms and charges in the half-reactions.

a. In this case, the sulfur atoms are balanced on both sides of the equation and the manganese atoms are balanced on both sides of the equation. Recall that in acid solution, H_2O molecules are available in abundance and can be used to balance oxygen atoms in the half-reactions.

$$SO_2 + 2H_2O \rightarrow SO_4^{2-} + 2e^- \text{ (oxidation)}$$

$$MnO_4^- + 5e^- \rightarrow Mn^{2+} + 4H_2O \text{ (reduction)}$$

b. Recall also that in acid solution, H^+ ions are readily available and can be used to balance the charge in the half-reactions.

$$SO_2 + 2H_2O \rightarrow SO_4^{2-} + 2e^- + 4H^+ \text{ (oxidation)}$$

$$MnO_4^- + 5e^- + 8H^+ \rightarrow Mn^{2+} + 4H_2O \text{ (reduction)}$$

Step 4. Adjust the coefficients so that the number of electrons lost in oxidation (2) equals the number of electrons gained in reduction (5). In this case, the least common multiple of 2 and 5 is 10. Cross-multiplying gives the balanced oxidation and reduction half-reactions.

$$5SO_2 + 10H_2O \rightarrow 5SO_4^{2-} + 20H^+ + 10e^- \text{ (oxidation)}$$

$$2MnO_4^- + 16H^+ + 10e^- \rightarrow 2Mn^{2+} + 8H_2O \text{ (reduction)}$$

Step 5. Add the balanced half-reactions and simplify by canceling or reducing like terms on both sides of the equation.

$$5SO_2 + 2H_2O + 2MnO_4^- \rightarrow 5SO_4^{2-} + 4H^+ + 2Mn^{2+}$$

Finally, return spectator ions (K^+) and restore the state descriptions. In this case, two K^+ ions go with the two MnO_4^- anions on the left side of the equation. Therefore, you can add only two K^+ ions to the right side, which will form K_2SO_4. The remaining SO_4^{2-} ions recombine with Mn^{2+} and H^+ to form two molecules of $MnSO_4$ and H_2SO_4, respectively.

$$5SO_2(g) + 2H_2O(l) + 2KMnO_4(aq) \longrightarrow K_2SO_4(aq) + 2H_2SO_4(aq) + 2MnSO_4(aq)$$

3. Evaluate the Answer

A review of the balanced equation indicates that the number of atoms of each element is equal on both sides of the equation. No subscripts have been changed.

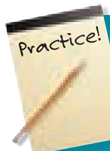
PRACTICE PROBLEMS

Use the half-reaction method to balance the following redox equations. Begin with step 2 of Example Problem 20-5, and leave the balanced equation in ionic form.

24. $Cr_2O_7^{2-}(aq) + I^-(aq) \longrightarrow Cr^{3+}(aq) + I_2(s)$ (in acid solution)

25. $Mn^{2+}(aq) + BiO_3^-(aq) \longrightarrow MnO_4^-(aq) + Bi^{2+}(aq)$ (in acid solution)

26. $N_2O(g) + ClO^-(aq) \longrightarrow NO_2^-(aq) + Cl^-(aq)$ (in basic solution). Hint: Add O and H in the form of OH^- ions and H_2O molecules.



For more practice solving problems using the half-reaction method, go to **Supplemental Practice Problems** in Appendix A.

You have learned two methods by which redox equations can be balanced. Both methods will provide the same results; however, the half-reaction method may be more useful to your study of electrochemistry.

Section 20.3 Assessment

27. Explain why an oxidation process must always accompany a reduction process.

28. Write the oxidation and reduction half-reactions for this redox equation.

$$Pb(s) + Pd(NO_3)_2(aq) \longrightarrow Pb(NO_3)_2(aq) + Pd(s)$$

29. Use the half-reaction method to balance this redox equation. Begin with step 2 of Example Problem 20-5, and leave the balanced equation in ionic form.

$$AsO_4^{3-}(aq) + Zn(s) \longrightarrow AsH_3(g) + Zn^{2+}(aq)$$
(in acid solution)

30. **Thinking Critically** The oxidation half-reaction of a redox reaction is $Sn^{2+} \longrightarrow Sn^{4+} + 2e^-$ and the reduction half-reaction is $Au^{3+} + 3e^- \longrightarrow Au$. What minimum numbers of tin(II) ions and gold(III) ions would have to react in order to have no electrons left over?

31. **Calculating** The concentration of thallium(I) ions in solution may be determined by oxidizing to thallium(III) ions with an aqueous solution of potassium permanganate ($KMnO_4$) under acidic conditions. Suppose that a 100.00 mL sample of a solution of unknown Tl^+ concentration is titrated to the endpoint with 28.23 mL of a 0.0560*M* solution of potassium permanganate. What is the concentration of Tl^+ ions in the sample? You must first balance the redox equation for the reaction to determine its stoichiometry.

$$Tl^+(aq) + MnO_4^-(aq) \longrightarrow Tl^{3+}(aq) + Mn^{2+}(aq)$$

Redox Reactions

In Section 20.2, a redox reaction involving copper and nitric acid is discussed. This reaction is balanced by a method called the oxidation-number method. In this lab, you will carry out this reaction, along with another redox reaction that involves a common household substance. You will practice balancing various redox reactions using both the oxidation-number method (from Section 20.2) and the half-reaction method (from Section 20.3).

Problem

What are some examples of redox reactions and how can the equations describing them be balanced?

Objectives

- **Observe** various redox reactions.
- **Balance** redox reactions using the oxidation-number method.
- **Balance** redox reactions using the half-reaction method.

Materials

copper metal
6*M* nitric acid
evaporating dish
forceps
distilled water
dropper pipette
spoon
household ammonia
crystal drain cleaner
thermometer
250-mL beaker

Safety Precautions

- **The reaction of copper with nitric acid should be done in a ventilation hood. Do not breathe the fumes from this reaction.**
- **Nitric acid and ammonia can cause burns. Avoid contact with skin and eyes.**

Pre-Lab

1. Read the entire **CHEMLAB**.
2. Prepare all written materials that you will take into the laboratory. Be sure to include safety precautions, procedure notes, and a data table.

Data Table	
Step 1	
Step 2	
Step 3	
Step 4	
Step 5	

3. Review what a redox reaction is.
4. Read the label of the crystal drain cleaner package. Understand that the compound is solid sodium hydroxide that contains aluminum. When the material is added to water, sodium hydroxide dissolves rapidly, producing heat. Aluminum reacts with water in the basic solution to produce $Al(OH)_4^-$ ions and hydrogen gas. Is aluminum oxidized or reduced in the reaction? Is hydrogen oxidized or reduced in the reaction? Explain your answers.

Procedure

1. In a ventilation hood, place a piece of copper metal in a clean, dry evaporating dish. Add enough 6*M* nitric acid to cover the metal. **CAUTION:** *Nitric acid can cause burns. The reaction of nitric acid with copper generates dangerous fumes. Use a ventilation hood.* Observe what happens and record your observations in the data table.
2. Pour about 2 mL of the solution from the evaporating dish into a test tube that contains about 2 mL of distilled water. Add ammonia until a change occurs. Record your observation in the data table.
3. Add about 50 mL of tap water to a 250-mL beaker. Use a thermometer to measure the temperature of the water. Record your observations in the data table.

4. Pour approximately 1 cm^3 of dry drain cleaner onto a watchglass. **CAUTION:** *Drain cleaner is caustic and will burn skin. Use forceps to move the crystals and observe their composition.* Record your observations in the data table.

5. Carefully pour about one-half spoonful of the crystals into the water in the beaker. As the crystals react with the water, watch the thermometer in the water for a few minutes and record in the data table the highest temperature reached and any other observations you make.

Cleanup and Disposal

1. After step 1 is completed, use forceps to remove any excess pieces of copper metal. Rinse the copper metal with tap water and dispose of the metal as your teacher instructs.
2. After step 2 is finished, pour the solution down the drain and flush with a lot of water.
3. After step 5 is finished, pour the solution down the drain and flush with a lot of water.

Analyze and Conclude

1. **Applying Concepts** The reaction between copper and nitric acid is discussed in Section 20.2. Write the half-reaction for the substance that is oxidized.
2. **Applying Concepts** Write the half-reaction for the substance that is reduced.
3. **Thinking Critically** In step 2, a deep blue copper–ammonia complex is formed according to the following reaction.

   $Cu^{2+} + NH_3 \longrightarrow Cu(NH_3)_4^{\ 2+}$

   Is this a redox reaction? Why or why not?
4. **Using Numbers** The following side reaction occurs from the reaction of copper with nitric acid.

   $Cu + HNO_3 \longrightarrow Cu(NO_3)_2 + NO + H_2O$

   Balance this redox reaction using both the oxidation-number method and the half-reaction method.
5. **Using Numbers** Write and balance the redox reaction of sodium hydroxide with aluminum and water.
6. **Error Analysis** Give possible reasons why you might not have been able to balance the equation for the redox reaction you performed in this experiment.

Real-World Chemistry

1. Using your observations in this lab, how do drain cleaning crystals remove clogs?
2. Ammonia and bleach are two common household chemicals that should never be mixed. One product of this reaction is chloramine, a poisonous, volatile compound. The reaction is as follows.

   $NH_3 + ClO^- \longrightarrow NH_2Cl + OH^-$

   What is the balanced redox reaction?
3. One type of breathalyzer detects whether ethanol is in the breath of a person. Ethanol is oxidized to acetaldehyde by dichromate ions in acidic solution. The dichromate ion in solution is orange, while the Cr^{3+} aqueous ion is green. The appearance of a green color in the breathalyzer test shows that the breath exceeds the legal limit of alcohol. The equation is

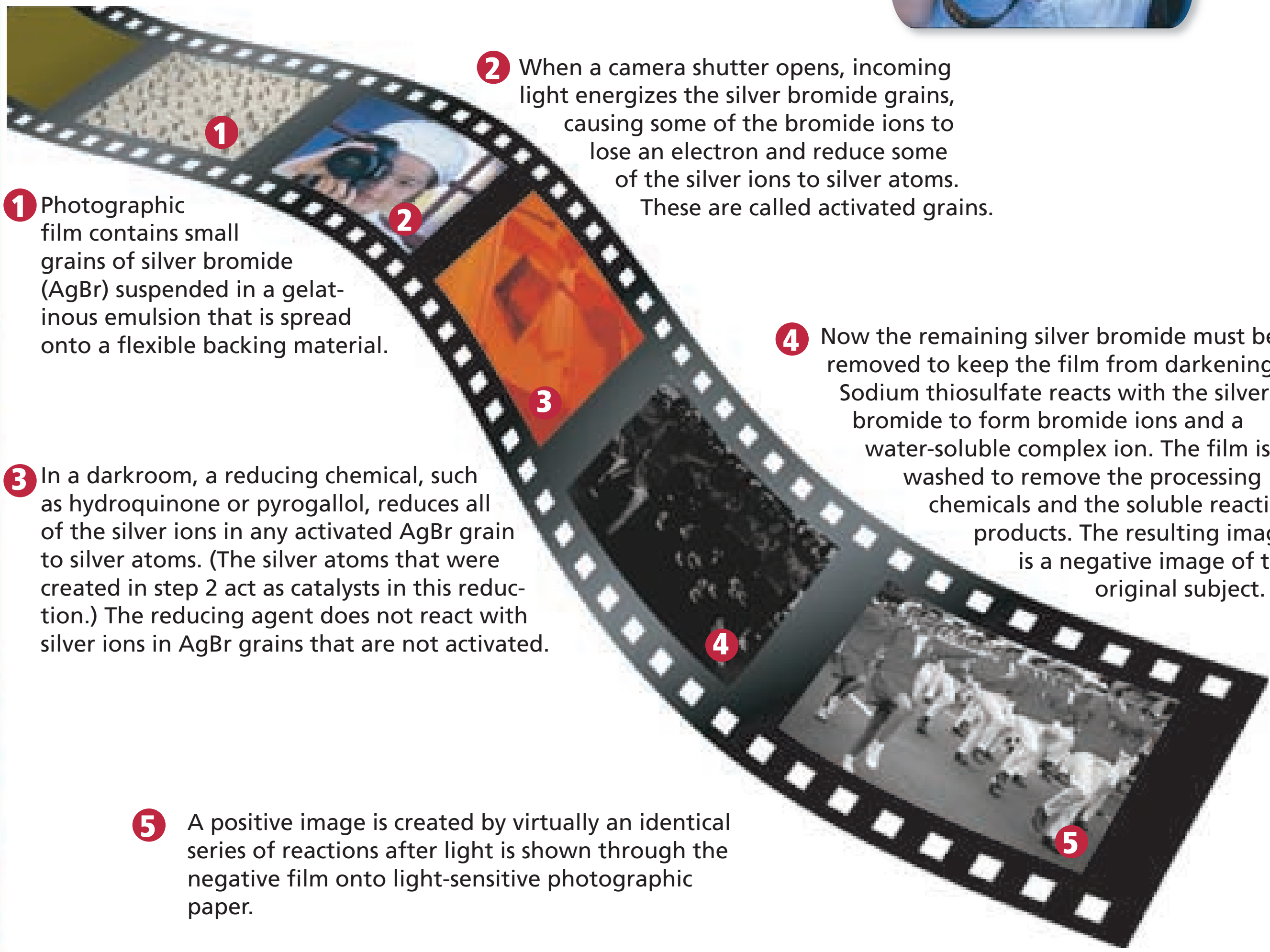
   $H^+ + Cr_2O_7^{\ 2-} + C_2H_5OH \longrightarrow Cr^{3+} + C_2H_4O + H_2O$

   Balance this redox reaction.
4. Diluted hydrochloric acid can be used to remove limestone (calcium carbonate) surrounding phosphate and silicate fossils. The reaction produces carbon dioxide, water, and aqueous calcium chloride. Write the balanced chemical equation. Is it a redox reaction? Explain.

How It Works

Photographic Film

Black and white photography is a popular art form. The spectacular images that are captured on and printed from black and white film are the product of a series of redox reactions. The first redox reaction captures the image on the film inside the camera. The second is a reaction to produce a negative image of the exposed film. The third redox reaction creates the positive print from the negative film image.



1 Photographic film contains small grains of silver bromide (AgBr) suspended in a gelatinous emulsion that is spread onto a flexible backing material.

2 When a camera shutter opens, incoming light energizes the silver bromide grains, causing some of the bromide ions to lose an electron and reduce some of the silver ions to silver atoms. These are called activated grains.

3 In a darkroom, a reducing chemical, such as hydroquinone or pyrogallol, reduces all of the silver ions in any activated AgBr grain to silver atoms. (The silver atoms that were created in step 2 act as catalysts in this reduction.) The reducing agent does not react with silver ions in AgBr grains that are not activated.

4 Now the remaining silver bromide must be removed to keep the film from darkening. Sodium thiosulfate reacts with the silver bromide to form bromide ions and a water-soluble complex ion. The film is washed to remove the processing chemicals and the soluble reaction products. The resulting image is a negative image of the original subject.

5 A positive image is created by virtually an identical series of reactions after light is shown through the negative film onto light-sensitive photographic paper.

Thinking Critically

1. **Predicting** Time, temperature, and solution concentration are all important factors in the developing process. What would be the consequence of leaving the film too long in the hydroquinone developer? *(Hint: the developer is a reducing agent.)*

2. **Hypothesizing** Color photographs are composed of dyes formed during the development of the silver image. The silver metal must be removed because it is opaque. Use the redox principle to explain how the silver could be removed from the color photo.

CONTENTS

CHAPTER 20 STUDY GUIDE

Summary

20.1 Oxidation and Reduction

- An oxidation–reduction (redox) reaction is any chemical reaction in which electrons are transferred from one atom to another. Most chemical reactions other than double replacement reactions are oxidation–reduction reactions.
- Oxidation occurs when an atom loses electrons. Reduction occurs when an atom gains electrons.
- Oxidation increases an atom's oxidation number and reduction decreases an atom's oxidation number.
- Oxidation and reduction must always occur together because the reduction process requires a supply of electrons available only from the oxidation process. At the same time, the oxidation process must have a receiver for lost electrons.
- The species that is oxidized in a redox reaction is the reducing agent and loses electrons. The species that is reduced is the oxidizing agent and gains electrons.
- The oxidation number of an element in a compound can be determined by applying a series of rules to the atoms in the compound.

Summary of Oxidation and Reduction Processes

Process	Oxidation	Reduction
Examples	$Na \rightarrow Na^+ + e^-$ $Fe^{2+} \rightarrow Fe^{3+} + e^-$	$Cl_2 + 2e^- \rightarrow Cl^-$ $Sn^{4+} + 2e^- \rightarrow Sn^{2+}$
Electron transfer	Atom loses electrons	Atom gains electrons
Change in oxidation number	Increases	Decreases
Function	Reducing agent	Oxidizing agent

20.2 Balancing Redox Equations

- Many simple redox equations may be balanced by inspection. The oxidation-number method can be used to balance more difficult reactions.
- Redox equations are often balanced by examining the change in oxidation number that occurs in the oxidized and reduced species. The following principle is then applied.

  e^- lost in oxidation = e^- gained in reduction

  Therefore, the total increase in oxidation numbers equals the total decrease in oxidation numbers.
- Redox reactions that take place in acidic solution often use water molecules and hydrogen ions in the process. Redox reactions that take place in basic solution often use water molecules and hydroxide ions in the process. Therefore, it is appropriate to add these species to the equation in order to balance the numbers of oxygen and/or hydrogen atoms.
- Redox reactions involving ionic species may be represented by net ionic equations, leaving out spectator ions.

20.3 Half-Reactions

- The oxidation and reduction processes of a redox reaction can be represented as half-reactions.
- An oxidation half-reaction shows the number of electrons lost when a species is oxidized. A reduction half-reaction shows the number of electrons gained when a species is reduced.
- The fact that an oxidation process must always be coupled with a reduction process provides a basis for using half-reactions to balance redox equations.

Vocabulary

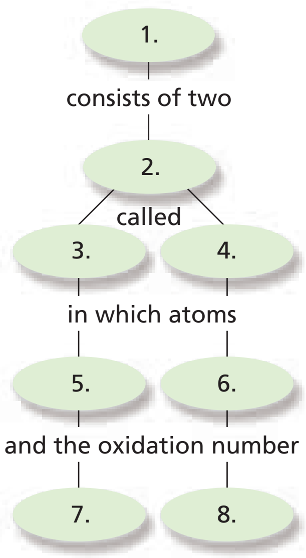
- half-reaction (p. 651)
- oxidation (p. 637)
- oxidation-number method (p. 644)
- oxidation–reduction reaction (p. 636)
- oxidizing agent (p. 638)
- redox reaction (p. 636)
- reducing agent (p. 638)
- reduction (p. 637)
- species (p. 650)

CHAPTER 20 ASSESSMENT

Go to the Chemistry Web site at **science.glencoe.com** *or use the* **Chemistry CD-ROM** *for additional Chapter 20 Assessment.*

Concept Mapping

32. Complete the concept map using the following terms: decreases, half-reactions, gain electrons, reduction, lose electrons, redox reaction, oxidation, increases.



Mastering Concepts

33. What is the main characteristic of oxidation–reduction reactions? (20.1)

34. In terms of electrons, what happens when an atom is oxidized? When an atom is reduced? (20.1)

35. What is the oxidation number of alkaline earth metals in their compounds? Of alkali metals? (20.1)

36. How does the oxidation number change when an element is oxidized? When it is reduced? (20.1)

37. Explain why oxidation and reduction must occur simultaneously. (20.1)

38. Identify the oxidizing agent and the reducing agent in the following equation. Explain your answer. (20.2)

$Fe(s) + Ag^{+}(aq) \longrightarrow Fe^{2+}(aq) + Ag(s)$

39. How can you tell that the redox equation in question 38 is not balanced? (20.2)

40. How does the change in oxidation number in an oxidation process relate to the number of electrons lost? How does the change in oxidation number in a reduction process relate to the number of electrons gained? (20.2)

41. Before you attempt to balance the equation for a redox reaction, why do you need to know whether the reaction takes place in acidic or basic solution? (20.3)

42. Does the following equation represent a reduction or an oxidation process? Explain your answer. (20.3)

$Zn^{2+} + 2e^{-} \longrightarrow Zn$

43. What term is used for the type of reaction represented in question 42? (20.3)

Mastering Problems

Oxidation and Reduction (20.1)

44. Identify the species oxidized and the species reduced in each of these redox equations.

a. $3Br_2 + 2Ga \longrightarrow 2GaBr_3$
b. $HCl + Zn \longrightarrow ZnCl_2 + H_2$
c. $Mg + N_2 \longrightarrow Mg_3N_2$

45. Identify the oxidizing agent and the reducing agent in each of these redox equations.

a. $H_2S + Cl_2 \longrightarrow 2HCl + S$
b. $N_2 + 3H_2 \longrightarrow 2NH_3$
c. $2Na + I_2 \longrightarrow 2NaI$

46. Identify each of these half-reactions as either oxidation or reduction.

a. $Al \longrightarrow Al^{3+} + 3e^{-}$
b. $NO_2 \longrightarrow NO_3^{-} + e^{-}$
c. $Cu^{2+} + e^{-} \longrightarrow Cu^{+}$

47. Determine the oxidation number of the bold element in these substances and ions.

a. $Ca\mathbf{Cr}O_4$
b. $NaH\mathbf{S}O_4$
c. $\mathbf{N}O_2^{-}$
d. $\mathbf{Br}O_3^{-}$

48. Determine the net change of oxidation number of each of the elements in these redox equations.

a. $C + O_2 \longrightarrow CO_2$
b. $Cl_2 + ZnI_2 \longrightarrow ZnCl_2 + I_2$
c. $CdO + CO \longrightarrow Cd + CO_2$

49. Which of these equations does **not** represent a redox reaction? Explain your answer.

a. $LiOH + HNO_3 \longrightarrow LiNO_3 + H_2O$
b. $Ag + S \longrightarrow Ag_2S$
c. $MgI_2 + Br_2 \longrightarrow MgBr_2 + I_2$

50. Identify each of these half-reactions as either oxidation or reduction.

a. $Fe^{3+} \longrightarrow Fe^{2+}$
b. $HPO_3^{2-} \longrightarrow HPO_4^{2-}$
c. $BCl_3 \longrightarrow B_2Cl_4$

51. Determine the oxidation number of nitrogen in each of these molecules or ions.

a. NH_3 **c.** N_2H_4 **e.** N_2O
b. KCN **d.** NO_3^- **f.** NF_3

52. Determine the oxidation number of each element in these compounds or ions.

a. $FeCr_2O_4$ (iron(II) chromite)
b. $Au_2(SeO_4)_3$ (gold(III) selenate)
c. $Ni(CN)_2$ (nickel(II) cyanide)

53. Explain how the sulfite ion (SO_3^{2-}) differs from sulfur trioxide (SO_3).

Balancing Redox Equations (20.2)

54. Use the oxidation-number method to balance these redox equations.

a. $CO + I_2O_5 \longrightarrow I_2 + CO_2$
b. $Cl_2 + NaOH \longrightarrow NaCl + HOCl$
c. $SO_2 + Br_2 + H_2O \longrightarrow HBr + H_2SO_4$
d. $HBrO_3 \longrightarrow Br_2 + H_2O + O_2$

55. Use the oxidation-number method to balance the following ionic redox equations.

a. $Al + I_2 \longrightarrow Al^{3+} + I^-$
b. $MnO_2 + Br^- \longrightarrow Mn^{2+} + Br_2$ (in acid solution)
c. $Cu + NO_3^- \longrightarrow Cu^{2+} + NO$ (in acid solution)
d. $Zn + NO_3^- \longrightarrow Zn^{2+} + NO_2$ (in acid solution)

56. Use the oxidation-number method to balance these redox equations.

a. $PbS + O_2 \longrightarrow PbO + SO_2$
b. $NaWO_3 + NaOH + O_2 \longrightarrow Na_2WO_4 + H_2O$
c. $NH_3 + CuO \longrightarrow Cu + N_2 + H_2O$
d. $Al_2O_3 + C + Cl_2 \longrightarrow AlCl_3 + CO$

57. Use the oxidation-number method to balance these ionic redox equations.

a. $MoCl_5 + S^{2-} \longrightarrow MoS_2 + Cl^- + S$
b. $Al + OH^- + H_2O \longrightarrow H_2 + AlO_2^-$
c. $TiCl_6^{2-} + Zn \longrightarrow Ti^{3+} + Cl^- + Zn^{2+}$

Half-Reactions (20.3)

58. Write the oxidation and reduction half-reactions represented in each of these redox equations. Write the half-reactions in net ionic form if they occur in aqueous solution.

a. $PbO(s) + NH_3(g) \longrightarrow N_2(g) + H_2O(l) + Pb(s)$
b. $I_2(s) + Na_2S_2O_3(aq) \longrightarrow Na_2S_2O_4(aq) + NaI(aq)$
c. $Sn(s) + 2HCl(aq) \longrightarrow SnCl_2(aq) + H_2(g)$

59. Use the half-reaction method to balance these equations. Add water molecules and hydrogen ions (in acid solutions) or hydroxide ions (in basic solutions) as needed. Keep balanced equations in net ionic form.

a. $Cl^-(aq) + NO_3^-(aq) \longrightarrow ClO^-(aq) + NO(g)$ (in acid solution)
b. $IO_3^-(aq) + Br^-(aq) \longrightarrow Br_2(l) + IBr(s)$ (in acid solution)
c. $I_2(s) + Na_2S_2O_3(aq) \longrightarrow Na_2S_2O_4(aq) + NaI(aq)$ (in acid solution)

60. Use the half-reaction method to balance these equations for redox reactions. Add water molecules and hydrogen ions (in acid solutions) or hydroxide ions (in basic solutions) as needed.

a. $NH_3(g) + NO_2(g) \longrightarrow N_2(g) + H_2O(l)$
b. $Mn^{2+}(aq) + BiO_3^-(aq) \longrightarrow Bi^{2+}(aq) + MnO_4^-(aq)$ (in acid solution)
c. $Br_2 \longrightarrow Br^- + BrO_3^-$ (in basic solution)

61. Balance the following redox chemical equation. Rewrite the equation in full ionic form, then derive the net ionic equation and balance by the half-reaction method. Give the final answer as it is shown below but with the balancing coefficients.

$KMnO_4(aq) + FeSO_4(aq) + H_2SO_4(aq) \longrightarrow$
$Fe_2(SO_4)_3(aq) + MnSO_4(aq) + K_2SO_4(aq) + H_2O(l)$

62. Balance this equation in the same manner as in question 61 above.

$HNO_3(aq) + K_2Cr_2O_4(aq) + Fe(NO_3)_2(aq) \longrightarrow$
$KNO_3(aq) + Fe(NO_3)_3(aq) + Cr(NO_3)_2(aq) + H_2O(l)$

Mixed Review

Sharpen your problem-solving skills by answering the following.

63. Determine the oxidation number of the bold element in each of the following examples.

a. $\mathbf{O}F_2$ **c.** $\mathbf{Ru}O_4$
b. $\mathbf{U}O_2^{2+}$ **d.** $\mathbf{Fe_2}O_3$

64. Identify the reducing agents in these equations.

a. $4NH_3 + 5O_2 \longrightarrow 4NO + 6H_2O$
b. $Na_2SO_4 + 4C \longrightarrow Na_2S + 4CO$
c. $4IrF_5 + Ir \longrightarrow 5IrF_4$

65. Write a balanced ionic redox equation using the following pairs of redox half-reactions.

a. $Fe \rightarrow Fe^{2+} + 2e^-$
$Te^{2+} + 2e^- \rightarrow Te$

b. $IO_4^- + 2e^- \rightarrow IO_3^-$
$Al \rightarrow Al^{3+} + 3e^-$ (in acid solution)

c. $I_2 + 2e^- \rightarrow 2I^-$
$N_2O \rightarrow NO_3^- + 4e^-$ (in acid solution)

66. Balance these redox equations by any method.

a. $P + H_2O + HNO_3 \rightarrow H_3PO_4 + NO$
b. $KClO_3 + HCl \rightarrow Cl_2 + ClO_2 + H_2O + KCl$

67. Balance these ionic redox equations by any method.

a. $Sb^{3+} + MnO_4^- \rightarrow SbO_4^{3-} + Mn^{2+}$ in acid solution
b. $N_2O + ClO^- \rightarrow Cl^- + NO_2^-$ in basic solution

68. Balance these ionic redox equations by any method.

a. $Mg + Fe^{3+} \rightarrow Mg^{2+} + Fe$
b. $ClO_3^- + SO_2 \rightarrow Cl^- + SO_4^{2-}$ (in acid solution)

Thinking Critically

69. Applying Concepts The following equations show redox reactions that are sometimes used in the laboratory to generate pure nitrogen gas and pure dinitrogen monoxide gas (nitrous oxide, N_2O).

$NH_4NO_2(s) \rightarrow N_2(g) + 2H_2O(l)$
$NH_4NO_3(s) \rightarrow N_2O(g) + 2H_2O(l)$

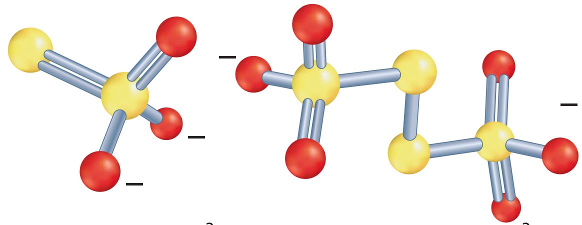
a. Determine the oxidation number of each element in the two equations and then make diagrams as in Example Problem 20-2 showing the changes in oxidation numbers that occur in each reaction.
b. Identify the atom that is oxidized and the atom that is reduced in each of the two reactions.
c. Identify the oxidizing and reducing agents in each of the two reactions.
d. Write a sentence telling how the electron transfer taking place in these two reactions differs from that taking place here.
$2AgNO_3 + Zn \rightarrow Zn(NO_3)_2 + 2Ag$

70. Comparing and Contrasting All redox reactions involve a transfer of electrons, but this transfer can take place between different types of atoms as illustrated by the two equations below. How does electron transfer in the first reaction differ from electron transfer in the second reaction?

$3Pb^{2+}(aq) + 2Cr(s) \rightarrow 2Cr^{3+}(aq) + 3Pb(s)$
$ClO_3^-(aq) + NO_2^-(aq) \rightarrow ClO_2^-(aq) + NO_3^-(aq)$

71. Using Numbers Examine the net ionic equation below for the reaction that occurs when the thiosulfate ion ($S_2O_3^{2-}$) is oxidized to the tetrathionate ion ($S_4O_6^{2-}$). Balance the equation using the half-reaction method. The structures of the two ions will help you to determine the oxidation numbers to use.

$S_2O_3^{2-} + I_2 \rightarrow I^- + S_4O_6^{2-}$ (in acid solution)



Thiosulfate ion ($S_2O_3^{2-}$) Tetrathionate ion ($S_4O_6^{2-}$)

72. Predicting Consider the fact that all of the following are stable compounds. What can you infer about the oxidation state of phosphorus in its compounds?

PH_3, PCl_3, P_2H_4, PCl_5, H_3PO_4, Na_3PO_3

Writing in Chemistry

73. Research the role of oxidation–reduction reactions in the manufacture of steel. Write a summary of your findings, including appropriate diagrams and equations representing the reactions.

74. Practice your technical writing skills by writing (in your own words) a procedure for cleaning tarnished silverware by a redox chemical process. Be sure to include background information describing the process as well as logical steps that would enable anyone to accomplish the task.

Cumulative Review

Refresh your understanding of previous chapters by answering the following.

75. How do the following characteristics apply to the electron configurations of transition metals? (Chapter 7)

a. Ions vary in charge.
b. Many elements have high melting points.
c. Many of their solids are colored.
d. Some elements are hard solids.

76. When iron(III) chloride reacts in an atmosphere of pure oxygen the following occurs:

$4FeCl_3(s) + 3O_2(g) \rightarrow 2Fe_2O_3(s) + 6Cl_2(g)$

If 45.0 g of iron(III) chloride reacts and 20.5 g of iron(III) oxide is recovered, determine the percent yield. (Chapter 12)

STANDARDIZED TEST PRACTICE
CHAPTER 20

Use these questions and the test-taking tip to prepare for your standardized test.

1. The reducing agent in a redox reaction is all of the following EXCEPT ______ .

a. the substance oxidized
b. the electron acceptor
c. the reducer of another substance
d. the electron donor

2. The oxidation numbers of the elements in $CuSO_4$ are ______ .

a. Cu = +2, S = +6, O = −2
b. Cu = +3, S = +5, O = −2
c. Cu = +2, S = +2, O = −1
d. Cu = +2, S = 0, O = −2

3. For the reaction $X + Y \longrightarrow XY$, the element that will be reduced is the one that is ______ .

a. more reactive
b. more massive
c. more electronegative
d. more radioactive

4. The net ionic reaction between iodine and lead(IV) oxide is shown below:

$I_2(s) + PbO_2(s) \longrightarrow IO_3^-(aq) + Pb^{2+}(aq)$

If the reaction takes place in acidic solution, the balanced equation is ______ .

a. $I_2(s) + 5PbO_2(s) + 4H_2O(l) \longrightarrow 2IO_3^-(aq) + 5Pb^{2+}(aq) + 8OH^-(aq)$
b. $I_2(s) + 5PbO_2(s) + H^+(aq) \longrightarrow 2IO_3^-(aq) + 5Pb^{2+}(aq) + H_2O(l)$
c. $I_2(s) + 5PbO_2(s) + 4H_2O(l) \longrightarrow 2IO_3^-(aq) + 5Pb^{2+}(aq) + 8H^+(aq)$
d. $I_2(s) + 5PbO_2(s) + 8H^+(aq) \longrightarrow 2IO_3^-(aq) + 5Pb^{2+}(aq) + 4H_2O(l)$

5. The reaction between sodium iodide and chloride is shown below:

$2NaI(aq) + Cl_2(aq) \longrightarrow 2NaCl(aq) + I_2(aq)$

The oxidation state of Na remains unchanged because ______ .

a. Na^+ is a spectator ion
b. Na^+ cannot be reduced
c. Na is an uncombined element
d. Na^+ is a monatomic ion

6. The reaction between nickel and copper(II) chloride is shown below:

$Ni(s) + CuCl_2(aq) \longrightarrow Cu(s) + NiCl_2(aq)$

The half reactions for this redox reaction are ______ .

a. $Ni \longrightarrow Ni^{2+} + 2e^-$, $Cl_2 \longrightarrow 2Cl^- + 2e^-$
b. $Ni \longrightarrow Ni^+ + e^-$, $Cu^+ + e^- \longrightarrow Cu$
c. $Ni \longrightarrow Ni^{2+} + 2e^-$, $Cu^{2+} + 2e^- \longrightarrow Cu$
d. $Ni \longrightarrow Ni^{2+} + 2e^-$, $2Cu^+ + 2e^- \longrightarrow Cu$

Interpreting Tables Use the table to answer questions 7–9.

Data for Elements in the Redox Reaction $Zn + HNO_3 \rightarrow Zn(NO_3)_2 + NO_2 + H_2O$

Element	Oxidation Number	Complex ion of which element is a part
Zn	0	none
Zn in $Zn(NO_3)_2$	+2	none
H in HNO_3	+1	none
H in H_2O	?	none
N in HNO_3	?	NO_3^-
N in NO_2	+4	none
N in $Zn(NO_3)_2$	?	NO_3^-
O in HNO_3	−2	NO_3^-
O in NO_2	?	none
O in $Zn(NO_3)_2$	?	NO_3^-
O in H_2O	−2	none

7. Which of these elements forms a monatomic ion that is a spectator in the redox reaction?

a. Zn
b. O
c. N
d. H

8. The oxidation number of N in $Zn(NO_3)_2$ is ______ .

a. +3
b. +5
c. +1
d. +6

9. The element that is oxidized in this reaction is _____ .

a. Zn
b. O
c. N
d. H

THE PRINCETON REVIEW

TEST-TAKING TIP

Write It Down! Most tests ask you a large number of questions in a small amount of time. Write down your work wherever possible. Write out the half reactions for a redox problem, and make sure they add up. Do math on paper, not in your head. Underline and reread important facts in passages and diagrams—don't try to memorize them.